5P

Worksheet: Formula Mass and Moles

Calculate the molar mass of the following:

1. MgO
2. CO_2
3. $NaOH$
4. H_2SO_4
5. $Ba(OH)_2$
6. $ZnSO_4 \cdot 7H_2O$
7. $Na_2CO_3 \cdot 10H_2O$
8. fluorine
9. sodium phosphate
10. iron(III) oxide

List the diatomic molecules:
List the polyatomic molecules:

Fill in the blanks:
1 mole=________________atoms/molecules=______________g=____________L

Use the mole definition and dimensional analysis to solve the following:

a. grams to atoms
300 gXe=_________atoms Xe

b. grams to moles
500 g Cl_2=___________moles Cl_2

c. moles to atoms
2.6 moles Kr=____________atoms Kr

d. moles to grams
5.3 moles F_2=__________moles F_2

e. atoms to moles
3.01×10^{23} atoms Mg=____________moles Mg

f. atoms to grams
1.2×10^{24} atoms Ar=_____________grams Ar

g. molecules to moles
6.02×10^{23} molecules O_2=____________ moles O_2

h. molecules to grams
1.2×10^{24} molecules Br_2=____________ grams Br_2

Worksheet: Moles

Use the mole definition and dimensonal analysis to solve the following:

1. 1 mole Fe=________g Fe

2. 1 mole Ni=_________ g Ni

3. .42 moles Na=_________ g Na

4. 134.68 g Pb= __________ moles Pb

5. 90.8 g Ne=___________ moles Ne

6. 4.6 moles Ag= __________ g Ag

7. 356 g B=___________ moles B

8. Calculate the mass in grams of 5.0 moles of carbon

9. Calculate the mass in grams of 10.5 moles of oxygen

10. Calculate the number of moles in 800 g of calcium

11. Calculate the number of molesin 560 g of bromine

12. Calculate the number of atoms in 3.2 moles of magnesium

13. Calculate the number of moles of atoms in 3.01×10^{23} atoms of zinc

14. Calculate the number of atoms in 0.5 moles of phosphorus.**

15. Calculate the mass in grams of 5.00 moles of ZnO

16. Calculate the mass in grams of 6.00 moles of hydrogen sulfate

17. Calculate the number of moles in 400 grams of sodium hydroxide.

18. Calculate the mass of 192 moles of hydrogen chloride

19. Calculate the number of molecules in 3.0 moles of methan, CH_4

20. Calculate the number of grams in 6.02×10^{23} molecules of $C_{12}H_{22}O_{11}$

21. Calculate the number of sulfur atoms produced when you start with 350 grams of sulfur molecules. **

EMPIRICAL FORMULA OF A COMPOUND

	Percent	Grams in 100 g sample	Moles	Mole ratio	Empirical formula
C	40.0%	40.0 g	$\frac{40.0}{12.0} = 3.33$	$\frac{3.33}{3.33} = 1$	
H	6.71%	6.71 g	$\frac{6.71}{1.01} = 6.64$	$\frac{6.64}{3.33} = 2$	CH_2O CHOH
O	53.3%	53.3 g	$\frac{53.3}{16.0} = 3.33$	$\frac{3.33}{3.33} = 1$	

Worksheet: Empirical Formula

1. Calculate the empirical formula for a compound that is 63.6% nitrogen and 36.4% oxygen.

2. Calculate the empirical formula for a compound that is 92.3% carbon and 7.7% hydrogen.

3. Calculate the empirical formula for a compound that is 57.5% sodium, 40.0% oxygen, and 2.5% hydrogen.

4. A compound is found to contain 39.95% carbon, 6.69% hydrogen, and 53.36% oxygen. Calculate the simplest formula of the compound.

5. Calculate the empirical formula for a compound that is 75.8% arsenic and 24.2% oxygen.
***

6. When 1.35 grams of silver oxide are decomposed, there remains a silver residue of 1.26 grams. Calculate the simplest formula of silver oxide.***

Worksheet: Molecular Formula

1. Find the molecular formula for a compound with percent composition 85.6% carbon, 14.4% hydrogen, and a molecular mass of 42.1g.

2. What is the molecular formula of cyanuric chloride if the empirical formula is CClN and the molecular mass is 184.5 g.

3. What is the molecular formula fo a substance with empirical formula $TlC_2H_2O_3$ and molecular mass of 557 g.

4. Hydrogen peroxide is found by analysis to consist of 5.9% hydrogen and 94.1% oxygen. Its molecular mass is 34.0 g. What is its molecular formula?

5. By analysis, a compound is found to be 76% iodine and the rest is oxygen. Its molecular mass is 334 g. What is its molecular formula? ***

60

Worksheet: Percent Composition

1. Determine the percentage of sodium in sodium sulfate, Na_2SO_4.

2. Calculate the percentage composition of nitrogen in each of the following:
a. HNO_2

b. NH_4NO_3

3. Calculate the mass of the metal in each of the following:
a. 50.0 grams MgS

b. 80.0 grams $FeCO_3$

c. 200.0 grams $Mg_3(PO_4)_2$

4. In a laboratory experiment, barium chloride dihydrate, $BaCl_2 \cdot 2H_2O$, is heated to remove completely its water of hydration. Calculate:
a. the experimental percent of water in the hydrate
b. the experimental percent of barium chloride(anhydrous) in the hydrate
c. the theoretical percent of water in the hydrate
d. the percent error of water in the hydrate based on the data below obtained in lab.

empty crucible and cover	20.286 g
crucible, cover and contents before heating	21.673 g
crucible, cover, and contents after heating	21.461 g

e. determine the experimental mol:mol ratio between the anhydrous and the hydrate.

5. The mining industry often reports the concentration of metal in an ore in terms of the amount of oxide formed by the metal. Naturally, this figure does not represent the actual amount of metal present. If a rock ore sample is analyzed as containing 1.00% by mass of iron(III) oxide, what is the percent by mass of the iron in the rock? ***

6. A sample of brass contains by mass 28.0% zinc and 72.0% copper. How many grams of brass can be produced from 6.00 grams of copper? ***

7. A compound has 12.32 g of hydrogen, 7.82 grams of oxygen, and 10.20 g sodium. What is the percent composition of each element?

Worksheet: Everything!!
Molecular mass, moles, % comp, empirical formula

1. Determine the molecular mass of HNO_3

2. Determine the molecular mass of $Fe_4[Fe(CN)_6]_3$

3. How many moles are in 24.5 g of H_2SO_4?

4. How many moles are in 95.4 g of Cu?

5. How many grams are in 0.3 moles of oxygen?

6. How many grams are in 3.1 moles of NaOH?

7. Find the percent composition of hydrogen in H_3PO_4

8. Find the percentage of arsenic in C_2H_8AsB

9. Find the percentage of copper in $CuCO_3 \cdot Cu(OH)_2$

10. Calculate the formula of a compound, given that 55.85 g of iron combines with 32.06 g of sulfur.

Worksheet: Everything!!!
Molecular mass, moles, molecular formula, empirical formula, percent composition

1. What is the molecular mass of glycerol, $C_3H_5(OH)_3$?

2. What is the molecular mass of nickel(II) sulfate hexahydrate, $NiSO_4 \cdot 6H_2O$?

3. How many moles are in 500 g of sucrose, $C_{12}H_{22}O_{11}$?

4. How many moles are in 1250 g of liquid methanol, CH_3OH?

5. How many grams are in 8.78 moles of benzene, C_6H_6?

6. How many grams are in 3.26 moles of glucose, $C_6H_{12}O_6$?

7. Calculate the molecular formula of a compound whose molecular mass is 128 g. Its percentage composition is 93.7% carbon and the rest is hydrogen.

8. Calculate the empirical formula of a compound that is 41.4% strontium, 13.24% nitrogen, and the rest is oxygen.

9. Calculate the percentage of water in the following hydrate: $Na_2CO_3 \cdot 10H_2O$

10. In a 10 gram sample of chloromycetin, 4.088 grams is carbon, 0.375 grams is hydrogen, 0.867 grams is nitrogen, 2.476 grams is oxygen, and 2.194 grams is chlorine. Calculate its empirical formula.

Mole Relationship in a Chemical Reaction — 8

The Law of Conservation of Matter, which was accepted for many years as absolute, stated that, "matter can be neither created nor destroyed." In other words, the total mass (the measure of the quantity of matter present) of the reactants must always equal the total mass of the products.

The discovery and development of nuclear reactions, however, forced the "law" to be revised. It was found that in nuclear reactions, very small but significant amounts of mass were lost and energy gained. The total mass of the reactants was not equal to the total mass of the products. This discovery led chemists to question the validity of the law in relation to ordinary or non-nuclear type reactions which are by far the most common in the study of chemistry. It was found, however, that in ordinary chemical reactions, interconversion of matter and energy is so small that it cannot be measured by any chemical balance. The Law of Conservation of Matter is therefore assumed to be correct in these reactions.

Objectives

In this experiment, you will test the Law of Conservation of Matter by causing a reaction to occur with a given amount of reactant. You will then carefully determine the mass of one of the products. With these measurements, you will be able to compare the moles. From the balanced equation, you should be able to see some relationship between the moles of reactant and moles of product in a chemical reaction.

Equipment

evaporating dish
watch glass
wire gauze
dropper pipet

Procedure

1. Clean an evaporating dish and rinse it with distilled water. Then heat it to a high temperature to insure the removal of all moisture.
2. Cool the dish and measure its mass to the nearest 0.01 g.
3. With a spatula, add about 3 g of sodium hydrogen carbonate ($NaHCO_3$) to the evaporating dish and read the mass to the nearest 0.01 g. You should not attempt to measure exactly 3.00 g since it is only a reference point, but your measurement of mass must be to the nearest 0.01 g. Just being close to 3.00 g is good. For example: 2.88 g or 3.08 g.
4. Obtain about 6 mL of 6*M* hydrochloric acid in a clean test tube and gradually add the acid to the $NaHCO_3$ with a dropper pipet. Allow the drops to enter the lip of the evaporating dish so that they flow down the side gradually. (See Figure 8-1)
5. Continue adding the acid slowly until the reaction has stopped. Do not add more acid than is needed.

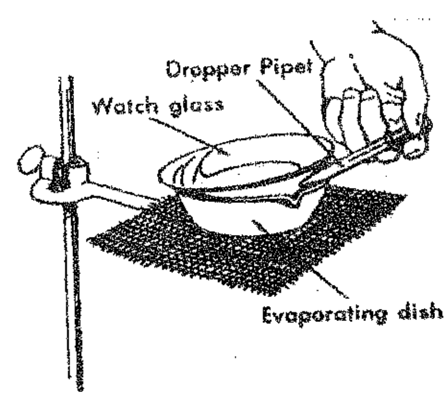

FIGURE 8-1. Method of adding acid slowly with a dropper pipet.

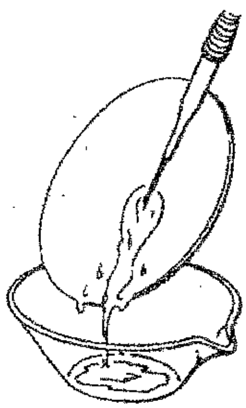

FIGURE 8-2. Rinsing the watch glass with distilled water.

6. Tilt the dish from side to side to make sure the liquid has reached all of the solid.
7. Remove the watch glass cover and rinse the underside of the watch glass with a very small amount of distilled water. Be careful to wash all material into the evaporating dish. (See Figure 8-2)
8. Heat the liquid in the evaporating dish over a water bath until it boils gently. Take care to avoid loss of liquid from boiling over.
9. Continue to dry the solid slowly until it appears to be completely dry.
10. Remove the dish from the water bath and allow it to cool, then measure and record the mass.
11. Reheat the dish and contents strongly, let it cool, and remeasure the mass. If this mass does not agree within 0.02 g with the mass in Step 10, reheat and remeasure the mass until a constant mass is achieved.
12. Continue to reheat until the last two mass measurements are within 0.02 g.

Data

Your data table should contain the following:

mass of empty evaporating dish	21.55 g
mass of evaporating dish + $NaHCO_3$	24.66 g
mass of $NaHCO_3$	3.11 g
mass of evaporating dish + NaCl	23.83 g
mass of NaCl	2.28 g

Calculations

1. One of the products was NaCl. The other two products were gases, one of which, when bubbled through limewater, causes the limewater to turn cloudy. Write the balanced equation for the reaction in this experiment.
2. Calculate the number of moles of $NaHCO_3$ used in this decomposition.
3. Calculate the number of moles of NaCl in the products.

Questions and Problems

1. From your balanced equation, what is the mole ratio between the $NaHCO_3$ and NaCl? How does it compare with the molar ratio from your experiment?
2. Suppose you had started with 3.25 moles of sodium hydrogen carbonate (sodium bicarbonate); how many moles of sodium chloride would you expect to be formed?
3. Does your experiment prove conclusively the Law of Conservation of Matter? Explain.

Data Table: Mole Relationship in a Chemical Reaction

Get it Signed!!!!

1. mass of watch glass ______g

2. mass of evaporating dish and watch glass ______g

3. mass of evaporating dish, watch glass and $NaHCO_3$ ______g

4. mass of $NaHCO_3$ ______g

5. mass of evaporating dish, watch glass and NaCl ______g ______g

6. mass of NaCl ______g

Chemistry Lab: Percent Composition of MgO

Purpose: To determine the % composition of Mg and O in MgO

Introduction: Chemists have found that when compounds are formed, elements always combine with one another in definite proportions or quantities by weight. For example, when water is formed, every 1 gram of hydrogen is combined with 8 grams of oxygen. This principle will be proven in lab by forming a compound of magnesium and oxygen and determine the percent composition.

Procedure:
1. Cut a piece of magnesium ribbon 30 cm long. Remove all the magnesium oxide from the piece by rubbing it with steel wool. Wipe clean with a paper towel.
2.Mass a crucible and cover to the nearest 0.01g. Record this mass in your data table.
3. Pinch the magnesium ribbon into a tiny wad and place it in the crucible. Mass the crucible, cover and sample to the nearest 0.01 g. Record this mass in your data table. Do NOT touch the magnesium or the crucible and cover with your hands after this step.
4. Place the crucible on a clay triangle and a ring stand and ring, only touching the crucible with your tongs. Remove the lid with your tongs. Heat the bottom of the crucible with a blue flame for 3 minutes. Now, increase the flame to an inner blue flame. **At the instant the magnesium ignites, put the cover on the crucible.**
5. After a brief period, raise the cover about an inch to admit air and close it again the instant the magnesium begins to burn.
6. Continue the process untill the sample no longer burns(reignites). When this happens, cover about 7/8 of the crucible with the lid and heat the bottom of the crucible strongly for about 5 more minutes.
7. Remove the flame. Place the crucible and cover on a wire gauze pad to cool. It will take about 5 minutes.
8. Mass the crucible and the cover and contents(powder) to the nearest 0.01g .**Only put the crucible and cover on the balance when it is cool** . Record the mass in your data table.
9. Put the contents(powder) on a paper towel. Add water to the powder. Do you detect an odor? What does it smell like? When you are finished, rinse the crucible and cover out with water and return to the bin and dispose of the paper towel in the trash.

Analysis:
1. Calculate the mass of magnesium.
2. Calculate the mass of oxygen combined with the magnesium.
3. Calculate the experimental % of magnesium and oxygen in your sample.
4. Write the correct formula for magnesium oxide based on the charges from the periodic table. Now, calculate the theoretical % of magnesium and oxygen in magnesium oxide using the masses off the periodic table.
5. Calculate a percent error for magnesium in magnesium oxide and a percent error for oxygen in magnesium oxide.
6. Now using your lab data, calculate the moles of magnesium and moles of oxygen in your sample
7. Using these mole values, find the simplest whole number mole:mole ratio between Mg andO. Use these values to state its experimental empirical formula.
8. If the molecular mass of magnesium oxide is found to be 40.3 g/mol, what is the molecular formula?
9. Define the Law of Definite Proportions(aka: Law of Definite Composition)
10. Based on your knowledge of why elements combine, can you explain why the proportions are always the same? Did your lab results prove the Law of Definite Proportions? Explain.

11. Could you use the % composition as a tool for identification of a compound? Explain.
12. Would % composition always identify the compound? (What about dimethyl ether, CH_3OCH_3 and ethanol, C_2H_5OH)
13. List 3 possible lab errors.

Fill in the Chart:

Substance	element or compound	meta or nonmetal	physical state solid, liquid or gas	color	combustibility yes or no?
magnesium					
oxygen					
magnesium oxide		XXXXXX			

	Symbol or formula
magnesium	
oxygen	
magnesium oxide	

Data Table: Percent Composition of MgO

Get it Signed!!!!!!

1. Mass of empty crucible and cover __________g **
2. Mass of crucible, cover and magnesium sample __________g **
3. Mass of magnesium sample __________g
4. Mass of crucible, cover and MgO __________g **
5. Mass of magnesium oxide, MgO __________g
6. Mass of oxygen that combined __________g
7. Percent oxygen in MgO (experimental) __________%
8. Percent oxygen in MgO (theoretical) __________%
9. Percent error of oxygen in MgO __________%

Lab: Formula of a Hydrate

Purpose: In this lab, you will determine the amount of water that a given hydrate contains. You will do this by heating the hydrate which will drive the water off. By knowing the mass of the hydrate and the mass of the anhydrous(without water) form, you can determine the number of water molecules(or % water in the hydrate) attached to each molecule of salt.

Procedure:
a. Obtain a crucible and cover and wash it out with water. You will not be able to get the crucible white, but do the best you can. We will assume that anything in the crucible before will also be there after.
b. Make sure the crucible is dry. You may have to heat it a little. Then mass it to the nearest 0.01g. Place this in the data table. After this step, you may only touch the crucible and cover with crucible tongs.
c. You will then receive a hydrate in your crucible about half full. Again, mass this and place this value in the data table. Don't forget to include the cover.
d. Place the crucible in a pipestem triange with the cover slightly ajar and heat with a blue flame for 3 minutes. Then change the flame to an inner blue flame and heat strongly for about 10 to 15 minutes.
e. Turn off the heat and place the crucible and cover on a wire gauze pad to cool. Cool for 5 minutes. Make sure you use crucible tongs and don't touch the crucible or cover.
f. Mass the crucible, cover and anhydrous(without water) to the nearest 0.01g and place on the data table.
g. When done, place the anhydrous on a paper towel. Put a drop of water back onto the anhydrous. Observe what happens. Record results. Then throw out paper towel with the anhydrous and clean the crucible with some water and dry out with a fresh paper towel.
h. Return the crucible and cover to the bin.
i. Put away ring stand, iron ring etc. BUT be very careful because they are still HOT!!!

Analysis Questions:
1. What is the purpose in heating the crucible and cover in the beginning of the procedure?
2. Why is it necessary to let the crucible cool before measuring the mass?
3. Why should the mass of the crucible be measured immediately after the crucible cools and not later?
4. What is a hydrate?
5. Does the hydrate look "wet?" Why not.
6. When you heated the hydrate, did you see any water come off? Why not?
7. Why is it important to know the amount of water in a hydrate when you are changing from grams to moles or moles to grams?
8. Calculate the formula of the hydrate by calculating a mole: mole ratio.
9. Compare your experimental and theoretical values. They are not exactly the same. Account for this by listing at least 3 possible lab errors.

Conclusion:
3 paragraphs
Summarize procedure
Tell me what you learned
How do you know you accomplished the purpose

Name:

Data Table: Formula of a Hydrate

Substance_____

Experimental Data Table

a. Mass of crucible and cover ________g

b. Mass of crucible, cover and hydrate ________g

c. Mass of hydrate ________g

d. Mass of crucible, cover and anhydrous ________g

e. Mass of anhydrous ________g

f. Experimental % water in hydrate ________%

Now, use the periodic table:

g. Mass of water molecules ________g

h. Mass of hydrate ________g

i. Theoretical % water in hydrate ________%

j. % error of water in hydrate ________%

Observations of hydrate before heating:

Observations of anhydrous after heating:

Observations of the anhydrous placed on a paper towel and adding a drop of water after the experiment is done:

60

Lab: Empirical Formula Determination

Introduction: A hydrate compound is one in which there are water molecules contained within the compound. Anhydrous means the water molecules are no lnoger present in the compound. Each lab group will be given one of the following hydrates with the theoretical formula given in parentheses.

barium chloride dihydrate: $BaCl_2 \cdot 2H_2O$

manganese sulfate monohydrate: $MnSO_4 \cdot H_20$

copper(II) sulfate pentahydrate: $CuSO_4 \cdot 5H_20$

magnesium sulfate heptahydrate: $MgSO_4 \cdot 7H_20$

You will attempt to verify the formula of the hydrate by determining the mole ratio of anhydrous salt to water molecules from your experimental data.

Procedure:
a. Obtain a crucible and cover and wash it out with water. You will not be able to get the crucible white, but do the best you can. We will assume that anything in the crucible before will also be there after.
b. Make sure the crucible is dry. You may have to heat it a little. Then mass it to the nearest 0.01g. Place this in the data table. After this step, you may only touch the crucible and cover with crucible tongs. Why?
c. You will then receive a hydrate in your crucible about half full. Again, mass this and place this value in the data table. Don't forget to include the cover.
d. Place the crucible in a pipestem triangle with the cover slightly ajar and heat with a blue flame for 3 minutes. Then change the flame to an inner blue flame and heat strongly for 10 to 15 minutes.
e. Turn off the heat and place the crucible and cover on a wire gauze pad to cool with your crucible tongs. Cool for 5 minutes. Make sure you do not touch the crucible and cover with your fingers..
f. Mass the crucible, cover and anhydrous(without water) to the nearest 0.01g and place on the data table.
g. When done, place the anhydrous substance on a paper towel. Put a drop of water back onto the anhydrous. Observe what happens. Record the results. Then throw out the paper towel with the anhydrous and clean the crucible with some water and dry out with a fresh paper towel.
h. Return the crucible and cover to the bin.
i. Put away ring stand, iron ring, etc. "BUT be very careful because they are still HOT!!!!

Analysis Questions:
1. What is the purpose in heating the crucible and cover in the beginning of the procedure?
2. Why is it necessary to let the crucible cool before measuring the mass?
3. Why should the mass of the crucible be measured immediately after the crucible cools and not later?
4. What is a hydrate?
5. Does the hydrate look "wet?"
6. When you heated the hydrate, did you see any water come off? Why not?
7. Why is it important to know the amount of water in a hydrate when you are changing from grams to moles or moles to grams?
8. Calculate the formula of the hydrate by calculating a mole: mole ratio.
9. Compare your experimental and theoretical values. They are not exactly the same. Account for this by listing at least 3 possible lab errors.

Name:

Data Table: Formula of a Hydrate

Substance_____

Experimental Data Table

a. Mass of crucible and cover ________g

b. Mass of crucible, cover and hydrate ________g

c. Mass of hydrate ________g

d. Mass of crucible, cover and anhydrous ________g

e. Mass of anhydrous ________g

f. Experimental % water in hydrate ________%

Now, use the periodic table:

g. Mass of water molecules ________g

h. Mass of hydrate ________g

i. Theoretical % water in hydrate ________%

j. % error of water in hydrate ________%

Observations of hydrate before heating:

Observations of anhydrous after heating:

Observations of the anhydrous placed on a paper towel and adding a drop of water after the experiment is done:

Lab: Formula of a Hydrate

Purpose: In this lab, you will determine the amount of water that a given hydrate contains. You will do this by heating the hydrate which will drive the water off. By knowing the mass of the hydrate and the mass of the anhydrous(without water) form, you can determine the number of water molecules(or % water in the hydrate) attached to each molecule of salt.

Procedure:

a. Obtain a crucible and cover and wash it out with water. You will not be able to get the crucible white, but do the best you can. We will assume that anything in the crucible before will also be there after.

b. Make sure the crucible is dry. You may have to heat it a little. Then mass it to the nearest 0.01g. Place this in the data table. After this step, you may only touch the crucible and cover with crucible tongs.

c. You will then receive a hydrate in your crucible about half full. Again, mass this and place this value in the data table. Don't forget to include the cover.

d. Place the crucible in a pipestem triange with the cover slightly ajar and heat with a blue flame for 3 minutes. Then change the flame to an inner blue flame and heat strongly for about 10 to 15 minutes.

e. Turn off the heat and place the crucible and cover on a wire gauze pad to cool. Cool for 5 minutes. Make sure you use crucible tongs and don't touch the crucible or cover.

f. Mass the crucible, cover and anhydrous(without water) to the nearest 0.01g and place on the data table.

g. When done, place the anhydrous on a paper towel. Put a drop of water back onto the anhydrous. Observe what happens. Record results. Then throw out paper towel with the anhydrous and clean the crucible with some water and dry out with a fresh paper towel.

h. Return the crucible and cover to the bin.

i. Put away ring stand, iron ring etc. BUT be very careful because they are still HOT!!!

Analysis Questions:

1. What is the purpose in heating the crucible and cover in the beginning of the procedure?
2. Why is it necessary to let the crucible cool before measuring the mass?
3. Why should the mass of the crucible be measured immediately after the crucible cools and not later?
4. What is a hydrate?
5. Does the hydrate look "wet?" Why not.
6. When you heated the hydrate, did you see any water come off? Why not?
7. Why is it important to know the amount of water in a hydrate when you are changing from grams to moles or moles to grams?
8. Calculate the formula of the hydrate by calculating a mole: mole ratio.
9. Compare your experimental and theoretical values. They are not exactly the same. Account for this by listing at least 3 possible lab errors.

Conclusion:

3 paragraphs

Summarize procedure

Tell me what you learned

How do you know you accomplished the purpose

Name:

Data Table: Formula of a Hydrate

Substance_____

Experimental Data Table

a. Mass of crucible and cover ________g

b. Mass of crucible, cover and hydrate ________g

c. Mass of hydrate ________g

d. Mass of crucible, cover and anhydrous ________g

e. Mass of anhydrous ________g

f. Experimental % water in hydrate ________%

Now, use the periodic table:

g. Mass of water molecules ________g

h. Mass of hydrate ________g

i. Theoretical % water in hydrate ________%

j. % error of water in hydrate ________%

Observations of hydrate before heating:

Observations of anhydrous after heating:

Observations of the anhydrous placed on a paper towel and adding a drop of water after the experiment is done:

Advanced Chemistry

Series 1 Number 6

Formula of a Hydrate

Many salts which have been crystallized from a water solution appear to be perfectly dry, yet when heated yield large quantities of water. The crystals change form, even color sometimes, as the water is driven off. This suggests that water was present as part of the crystal structure. Such compounds are called hydrates. The number of moles of water present per mole of anhydrous salt is usually some simple number. One example is the hydrate copper sulfate. Its blue crystals look and feel dry. Yet, each mole of hydrate contains five moles of water. Its formula is: $CuSO_4{\bullet}5H_2O$. The dot between the two parts does not mean multiplication. It means that water molecules are rather loosely attached to the other atoms. The molecular weight of this compound must be calculated taking the weight of the water into account. It would be 249.6 g/mol. In this experiment, you will be given an appropriate hydrate. You will determine the mass of water driven off by heating and the amount of anhydrous salt which remains. Your teacher will give you the formula of the anhydrous salt so that you can find the empirical formula of the hydrate.

Procedure:

1) Place a clean, dry crucible with cover in a triangle mounted on an iron ring. Leave the cover askew so that any water in the crucible can easily escape. Heat with a Bunsen burner flame for two or three minutes. Remove the flame.
2) When the crucible is cool enough to touch, transfer the crucible and cover to a balance and weigh to 0.01 g.
3) Put enough of the hydrate crystals in the crucible to fill it 1/4 full. (approx. 5 g). Weigh the crucible cover and contents.
4) Place the crucible, with its cover only slightly askew, on the triangle and heat them very gently so as to avoid spattering. Continue heating gently until the popping and sizzling have stopped. Gradually increase the flame until the crucible bottom is at most a dull red. Adjust the cover until there is a 0.5 cm opening between it and the crucible. Maintain this temperature for five minutes. Use a pair of tongs to transfer carefully the hot crucible and cover to a desicator.
5) Wait about 5 minutes, until the crucible is cool enough to touch. Transfer it and the cover to the balance and weigh them.
6) To make sure all the water is driven off, heat the crucible and its cover to a dull redness again. Carefully transfer them to a desicator. When cool, weigh again. If your results do not agree within 0.03 g, consult your teacher concerning further heating and weighing.

Calculations:

1) Calculate the mass of anhydrous salt.

2) Calculate the mass of water driven off.
3) Convert values in #1 and #2 to moles.
4) Calculate the ratio of moles H_2O to moles of anhydrous.

Moles water/lMoles anhydrous = X
5) Write the formula including mole H_2O.

Questions:

1) Look up and define a) efflorescent b) hydroscopic c) deliquescence.
2) What is the purpose of the desicator? What kind of substance is in to bottom of the desicator? (see #1).
3) List 5 examples of hydrates with the correct formula. What is the highest number of water molecules you can find in a hydrate.
4) Name the hydrates in #3.
5) Can you suggest reasons why the procedure used in this experiment might not be suitable for all hydrates?
6) List possible experimental errors and tell how each one effects your answer. (at least 3).

Ohemistry Lab #5 Percent Composition of Zinc Chloride

Part I Reaction of Zinc with HCl
Procedure:
1) This lab will be done on two days. The first day the reaction will be run, and on the second day, th;e dry product will be weighed.
2) Clean and weigh a dry 250 ml beaker on the centigram balance to the nearest 0.01g. Into this beaker you will place about 25 ml of 6 M Hydrochloric acid. The 6 M refers to the concentration of the acid, and will be written on the lable. CAUTION: this hydrochoric acid solution is strong. It will burn your hands if you have prolonged contact with it. If you get any on your hands, simply flush them with lots of water. In addition, it will put holes in clothing that it comes in contact with, so you must wear an apron.
3) When we weigh any chemical or substance on our balances, it must be weighed in a container or on paper so that the chemical does not touch the balance pan directly. You will obtain a piece of weighing paper and weigh it to the nearest 0.01 g. This paper is made so that chemicals do not stick to it. Place the balance about 2.5 grams heavier than the paper weighed. Add zinc metal to it until the balance begins to move. Once the balance arm begins to move, measure the mass using the weights to balance the zinc. Record the mass of the zinc and paper together.
4) The reaction between the hydrochloric acid and the zinc will produce hydrogen gas. Therefore, before you mix the two together, make sure you are away from any open flames. Place the zinc metal into the HCl and watch the reaction. Do not breathe the fumes.
5) After the zinc has completely reacted, place your beaker under the hood. Your teacher will place it in the oven overnight so that the water will be removed. (Be sure that you beaker has your lab drawer number on it so you will be able to identify it the next day.

Part II Second Day
1) Weigh your beaker that contains the zinc chloride.
2) Clean out your beaker with warm water. The zinc chloride will dissolve.

Discussion Questions
1) Complete the calculations for the percent zinc and percent chlorine in your sample as shown by your teacher in class.
2) Does the percent composition vary with the amount of zinc or hydrochloric acid used. Explain.
3) Write the Law of Definit Composition. Explain it in your own words.
4) What possible errors occured in this lab.

I Title.

II Purpose

III Data — get it signed

Day 1	1)	250 ml beaker	_______ g
	2)	weigh paper	_______ g
	3)	weigh paper + Zn	_______ g
	4)	Zn (~2.5g)	_______ g
Day 2	5)	beaker + $ZnCl_2$	_______ g
	6)	$ZnCl_2$	_______ g

IV Calculations

a) Exp % Zn in $ZnCl_2$.

$$\frac{\text{line 4}}{\text{line 6}} \times 100 =$$

Exp % Cl_2 in $ZnCl_2$

$$\frac{\text{line 6} - \text{line 4}}{\text{line 6}} \times 100 =$$

OR

100 − above calculation =

b) Theoretical % Zn in $ZnCl_2$ (Use The periodic Table)

$$\frac{\text{g Zn}}{\text{g } ZnCl_2} \times 100 =$$

Theoretical % Cl_2 in $ZnCl_2$

$$\frac{\text{g } Cl_2}{\text{g } ZnCl_2} \times 100 =$$

OR 100 − above calculation =

c) % error of Zn $= \left|\frac{\text{theor} - \text{exp}}{\text{theor}}\right| \times 100 =$

V ~~Analysis~~ Questions

~~1)~~ See above calculations
2) Yes OR No. Explain
3) State the law. Explain
4) State 3 possible errors

VI Conclusion

3 paragraphs.
- Summarize procedure
- Tell me what you learned
- How you know you accomplished the purpose.

Name: Date:

Accelerated Chem: Take-Home Quiz

Show ALL work for credit!

My Metal is_____________.

Write the formula and name for this metal carbonate.____________ ______________

Calculate its molecular mass. Is this the empirical formula? Explain why or why not?

Calculate the percent composition of the metal in the above compound.

Write a balanced decomposition reaction involving this metal carbonate.

Write a balanced composition reaction that produces your metal hydroxide.

Calculate the molecular mass of the metal hydroxide. Is this its empiricial formula? Explain why or why not.

Calculate the percent composition of the metal in the compound used above.

Write a balanced single replacement reaction involving your metal and some other metal compound.

Is your metal stronger or weaker than Hydrogen? Explain.

Write a balanced single replacement reaction involving your metal and HCl acid.

Write a double replacement reaction involving your metal.

Does this reaction occur? Explain why or why not.

Are there any ppt's in your products? Explain why or why not.

State 3 physical properties of your metal.

State 3 chemical properties of your metal.

State 3 primary uses of your metal or the metal within a compound.